T08D04 – (18.1) AB Calcs Notes

Name …….

1. 18.1.1 State the expression for the ionic product constant of water (K_w). (1)
   a. Provide an equilibrium equation for water and the ionic product constant equation for water:

   b. Pure water has the same concentration of what two ions in solution? What is it at 25°C?

2. 18.1.2 Deduce $[H^+(aq)]$ and $[OH^-(aq)]$ for water at different temperatures given K_w values. (3)
   a. What happens to the value of K_w as temperature changes? Explain, complete the table, and provide a rough graph of the effect of temperature on K_w:

Temperature	0°C	25°C	40°C	100°C
K_w (mol^2dm^{-6})				

   b. How does this affect the values of $[OH^-]$ and $[H^+]$ in solution?

   c. How does this affect the pH of solutions?

   d. Explain, using principles of energetics and equilibrium, why the dissociation of water is an endothermic process:

   e. At 60°C, the ionic product constant of water is 9.6×10^{-14} mol^2dm^{-6}. Calculate the pH of a neutral solution at this temperature

3. 18.1.3 Solve problems involving $[H^+(aq)]$, $[OH^-(aq)]$, pH and pOH. (3)
   a. This material was covered in SL Topic 08, for review please revisit those notes for details
   b. Calculate the pH of 0.01 mol dm^{-3} NaOH(aq):

   c. Calculate the pH of 0.01 mol dm^{-3} $Ba(OH)_2$

   d. Calculate the pH of 0.01 mol dm^{-3} sulfuric acid: *Hint: think about the strong/weak properties of sulfuric acid and it's conjugates as it's deprotonated*

   e. A solution of HNO_3(aq) contains 1.26g of the pure acid in every 100cm^3 of aqueous solution. Calculate the pH of the solution

   f. A solution of NaOH(aq) contains 0.40g in every 50cm^3 of aqueous solution. Calculate the pH of the solution

g. Briefly explain (no need to show all the work), using 1.0×10^{-8} mol dm^{-3} HCl(aq), why solutions of dilute acids can cause issues with pH calculations:

h. Although we will not use the following cases in IB chemistry, provide examples of an acid and a base that reach beyond the 0-14 pH scale range:

4. 18.1.4 State the equation for the reaction of any weak acid or weak base with water, and hence deduce the expressions for K_a and K_b. (1)

a. Deduce the following expressions:

	What is it?	How is it calculated	How would you get to pK_a, pK_b, or pK_w respectively?
Acid Ionization Constant (K_a)		*Chem Eqn:*	$K_a \rightarrow pK_a$
		Eq. Expression:	$pK_a \rightarrow K_a$
Base ionization Constant (K_b)		*Chem Eqn:*	$K_b \rightarrow pK_b$
		Eq. Expression:	$pK_b \rightarrow K_b$
Ionic Product Constant for Water (K_w)		*Chem Eqn:*	$K_w \rightarrow pK_w$
		Eq. Expression:	$pK_w \rightarrow K_w$

b. Complete the following table for strong and weak acids and bases, noting which will have high or low values of K_a, K_b, pK_a, pK_b, etc

	Strong Acid	Weak Acid	Strong Base	Weak Base
Values for the reactant	K_a: pK_a:	K_a: pK_a:	K_b: pK_b:	K_b: pK_b:
Values for its CONJUGATES	K_b: pK_b:	K_b: pK_b:	K_a: pK_a:	K_a: pK_a:

c. Show, using equilibrium equations, how the equilibrium constants for acids and bases, K_a and K_b respectively, change for strong and weak materials:

	Strong Acid	Weak Acid	Strong Base	Weak Base
Equilibrium Expression				

d. How can the pH of the following be calculated?

	Strong Acid	Weak Acid	Strong Base	Weak Base
How to Calculate pH				

e. We make a very important assumption in dealing with weak acid and base calculations, explain how **assumptions are for the weak:**

5. 18.1.5 Solve problems involving solutions of weak acids and bases using the expressions: $K_a \times K_b = K_w$, $pK_a + pK_b = pK_w$, $pH + pOH = pK_w$. (3)
   a. *Included in SL calcs and below in further calculations.*
6. 18.1.6 Identify the relative strengths of acids and bases using values of K_a, K_b, pK_a and pK_b. (2)
   a. What does I.C.E. stand for and how does the chart help us to organize our calculations?

b. Calculate the pH of a 1.00 mol dm^{-3} aqueous solution of hydrofluoric acid, HF(aq), given that the acid dissociation constant for HF is 7.2×10^{-4}

$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$

$$K_a = \frac{H+\ aq\ \ x[F-\ aq\ ]}{[HF\ aq\ ]}$$

	HF(aq) M	H^+(aq) M	F^-(aq) M
Initial			
Change			
Equilibrium			

$$K_a = \frac{x\ [x]}{[1.00-x]} = \frac{[x]^2}{1.00-x}$$

$$K_a = 7.2\text{x}10\text{-}4 = \frac{H+\ aq\ \ x[F-\ aq\ ]}{[HF\ aq\ ]} = \frac{x\ [x]}{[1.00-x]} = \boxed{\frac{[x]^2}{1.00-x}} \approx \boxed{\frac{[x]^2}{1.00}}$$

Calculate the pH value using both the quadratic approach and the assumption, and then compare the two:

If no assumption, solve quadratic:
$x^2 + (7.2\text{x}10^{-4})x - (7.2\text{x}10^{-4} \text{ x } 1.00) = 0$
$ax^2 + bx + c = 0$
$a = 1$
$b = 7.2\text{x}10^{-4}$
$c = -7.2\text{x}10^{-4}$

$$\frac{}{} = 2.65\text{x}10^{-2} = [H^+]$$

$-\log[H^+] = pH = \mathbf{1.57}$

For the assumption:
$x^2 = 7.2\text{x}10^{-4} \text{ x } 1.00$
$x = \sqrt{\ }$
$x = 2.68\text{x}10^{-2}$
$x = [H^+]$
$-\log[H^+] = pH$
$pH = \mathbf{1.57}$

c. Things to know about pK values:
   i. pK_a and pK_b numbers are usually positive w/o units
   ii. Stronger acids or bases with high values for K_a or K_b have lower values for pK_a and pK_b respectively
   iii. A change in one unit of pK_a or pK_b represents a 10-fold change in the value of K_a or K_b
   iv. pK_a and pK_b must be quoted at a specific temperature, generally 25°C for standard conditions and measurements
   v. High values of K_a for an acid will have low values of K_b for its conjugate base (make sense?)

$$K_a \bullet K_b = [H^+][OH^-] = K_w$$

d. Complete the following tables to explain how the K_a and pK_a and the K_b and pK_b are related:

i. For Acids:

Acid	Formula	K_a	pK_a
Methanoic			
Ethanoic			
Propanoic			

ii. For Bases:

Bases	Formula	K_b	pK_b
Ammonia			
Methylamine			
Ethylamine			

e. Calculate the K_b for a 0.100 mol dm^{-3} aqueous solution of methylamine, CH_3NH_2, at 25°C. Its pH is 11.80 at this temperature

f. A 0.20 mol dm^{-3} aqueous solution of ammonia has K_b of 1.8×10^{-5} at 25°C, what is the pH?